Sample Exercise 5.1 Describing and Calculating Energy Changes

A bowler lifts a 5.4-kg (12-lb) bowling ball from ground level to a height of 1.6 m (5.2 ft) and then drops it. **(a)** What happens to the potential energy of the ball as it is raised? **(b)** What quantity of work, in J, is used to raise the ball? **(c)** After the ball is dropped, it gains kinetic energy. If all the work done in part (b) has been converted to kinetic energy by the time the ball strikes the ground, what is the ball's speed just before it hits the ground? (Note: The force due to gravity is $F = m \times g$, where m is the mass of the object and g is the gravitational constant; $g = 9.8$ m/s^2.)

Solution

Analyze We need to relate the potential energy of the bowling ball to its position relative to the ground. We then need to establish the relationship between work and the change in the ball's potential energy. Finally, we need to connect the change in potential energy when the ball is dropped with the kinetic energy attained by the ball.

Plan We can calculate the work done in lifting the ball by using Equation 5.3: $w = F \times d$. The kinetic energy of the ball just before it hits the ground equals its initial potential energy. We can use the kinetic energy and Equation 5.1 to calculate the speed, v, just before impact.

Solve
(a) Because the ball is raised above the ground, its potential energy relative to the ground increases.
(b) The ball has a mass of 5.4 kg and is lifted 1.6 m. To calculate the work performed to raise the ball, we use Equation 5.3 and $F = m \times g$ for the force that is due to gravity:
$$w = F \times d = m \times g \times d = (5.4 \text{ kg})(9.8 \text{ m/s}^2)(1.6 \text{ m}) = 85 \text{ kg-m}^2/\text{s}^2 = 85 \text{ J}$$

Thus, the bowler has done 85 J of work to lift the ball to a height of 1.6 m.

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Sample Exercise 5.1 Describing and Calculating Energy Changes

Continued

(c) When the ball is dropped, its potential energy is converted to kinetic energy. We assume that the kinetic energy just before the ball hits the ground is equal to the work done in part (b), 85 J:

$$E_k = \tfrac{1}{2} mv^2 = 85 \text{ J} = 85 \text{ kg-m}^2/\text{s}^2$$

We can now solve this equation for v:

$$v^2 = \left(\frac{2E_k}{m}\right) = \left(\frac{2(85 \text{ kg-m}^2/\text{s}^2)}{5.4 \text{ kg}}\right) = 31.5 \text{ m}^2/\text{s}^2$$

$$v = \sqrt{31.5 \text{ m}^2/\text{s}^2} = 5.6 \text{ m/s}$$

Check Work must be done in (b) to increase the potential energy of the ball, which is in accord with our experience. The units are appropriate in (b) and (c). The work is in units of J and the speed in units of m/s. In (c) we carry an additional digit in the intermediate calculation involving the square root, but we report the final value to only two significant figures, as appropriate.

Comment A speed of 1 m/s is roughly 2 mph, so the bowling ball has a speed greater than 10 mph just before impact.
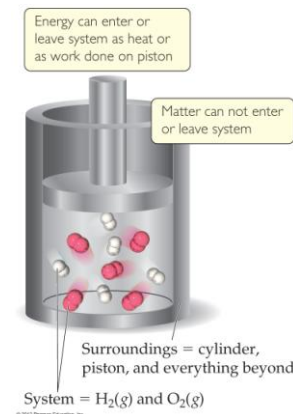
Practice Exercise

What is the kinetic energy, in J, of **(a)** an Ar atom moving at a speed of 650 m/s, **(b)** a mole of Ar atoms moving at 650 m/s? (*Hint*: 1 amu = 1.66×10^{-27} kg.)

Answer: **(a)** 1.4×10^{-20} J, **(b)** 8.4×103 J

Sample Exercise 5.2 Relating Heat and Work to Changes of Internal Energy

Gases A(*g*) and B(*g*) are confined in a cylinder-and-piston arrangement like that in Figure 5.4 and react to form a solid product C(*s*): A(*g*) + B(*g*) ⟶ C(*s*). As the reaction occurs, the system loses 1150 J of heat to the surroundings. The piston moves downward as the gases react to form a solid. As the volume of the gas decreases under the constant pressure of the atmosphere, the surroundings do 480 J of work on the system. What is the change in the internal energy of the system?

Energy can enter or leave system as heat or as work done on piston

Matter can not enter or leave system

Surroundings = cylinder, piston, and everything beyond
System = $H_2(g)$ and $O_2(g)$
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Solution

Analyze The question asks us to determine ΔE, given information about *q* and *w*.

Plan We first determine the signs of *q* and *w* (Table 5.1) and then use Equation 5.5, $\Delta E = q + w$, to calculate ΔE.

TABLE 5.1 • Sign Conventions for *q, w,* and ΔE

For *q*	+ means system *gains* heat	− means system *loses* heat
For *w*	+ means work done *on* system	− means work done *by* system
For ΔE	+ means *net gain* of energy by system	− means *net loss* of energy by system

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Sample Exercise 5.2 Relating Heat and Work to Changes of Internal Energy

Continued

Solve Heat is transferred from the system to the surroundings, and work is done on the system by the surroundings, so q is negative and w is positive: $q = -1150$ J and $w = 480$ kJ. Thus,

$$\Delta E = q + w = (-1150 \text{ J}) + (480 \text{ J}) = -670 \text{ J}$$

The negative value of ΔE tells us that a net quantity of 670 J of energy has been transferred from the system to the surroundings.

Comment You can think of this change as a decrease of 670 J in the net value of the system's energy bank account (hence, the negative sign); 1150 J is withdrawn in the form of heat while 480 J is deposited in the form of work. Notice that as the volume of the gases decreases, work is being done *on* the system *by* the surroundings, resulting in a deposit of energy.

Practice Exercise

Calculate the change in the internal energy for a process in which a system absorbs 140 J of heat from the surroundings and does 85 J of work on the surroundings.

Answer: +55 J

Sample Exercise 5.3 Determining the Sign of ΔH

Indicate the sign of the enthalpy change, ΔH, in these processes carried out under atmospheric pressure and indicate whether each process is endothermic or exothermic: **(a)** An ice cube melts; **(b)** 1 g of butane (C_4H_{10}) is combusted in sufficient oxygen to give complete combustion to CO_2 and H_2O.

Solution

Analyze Our goal is to determine whether ΔH is positive or negative for each process. Because each process occurs at constant pressure, the enthalpy change equals the quantity of heat absorbed or released, $\Delta H = q_P$.

Plan We must predict whether heat is absorbed or released by the system in each process. Processes in which heat is absorbed are endothermic and have a positive sign for ΔH; those in which heat is released are exothermic and have a negative sign for ΔH.

Solve In (a) the water that makes up the ice cube is the system. The ice cube absorbs heat from the surroundings as it melts, so ΔH is positive and the process is endothermic. In (b) the system is the 1 g of butane and the oxygen required to combust it. The combustion of butane in oxygen gives off heat, so ΔH is negative and the process is exothermic.

Practice Exercise

Molten gold poured into a mold solidifies at atmospheric pressure. With the gold defined as the system, is the solidification an exothermic or endothermic process?

Answer: In order to solidify, the gold must cool to below its melting temperature. It cools by transferring heat to its surroundings. The air around the sample would feel hot because heat is transferred to it from the molten gold, meaning the process is exothermic.

You may notice that solidification of a liquid is the reverse of the melting we analyzed in the exercise. As we will see, reversing the direction of a process changes the sign of the heat transferred.

Sample Exercise 5.4 Relating ΔH to Quantities of Reactants and Products

How much heat is released when 4.50 g of methane gas is burned in a constant-pressure system? (Use the information given in Equation 5.18.)

Solution

Analyze Our goal is to use a thermochemical equation to calculate the heat produced when a specific amount of methane gas is combusted. According to Equation 5.18, 890 kJ is released by the system when 1 mol CH_4 is burned at constant pressure.

Plan Equation 5.18 provides us with a stoichiometric conversion factor: (1 mol CH_4 = 890 kJ). Thus, we can convert moles of CH_4 to kJ of energy. First, however, we must convert grams of CH_4 to moles of CH_4. Thus, the conversion sequence is grams CH_4 (given) → moles CH_4 → kJ (unknown to be found).

Solve By adding the atomic weights of C and 4 H, we have 1 mol CH_4 = 16.0 CH_4. We can use the appropriate conversion factors to convert grams of CH_4 to moles of CH_4 to kilojoules:

$$\text{Heat} = (4.50 \text{ g } CH_4)\left(\frac{1 \text{ mol } CH_4}{16.0 \text{ g } CH_4}\right)\left(\frac{-890 \text{ kJ}}{1 \text{ mol } CH_4}\right) = -250 \text{ kJ}$$

The negative sign indicates that the system released 250 kJ into the surroundings.

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Sample Exercise 5.4 Relating ΔH to Quantities of Reactants and Products

Continued

Practice Exercise

Hydrogen peroxide can decompose to water and oxygen by the reaction

$$2\,H_2O_2(l) \longrightarrow 2\,H_2O(l) + O_2(g) \qquad \Delta H = -196\text{ kJ}$$
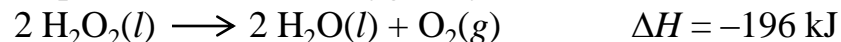
Calculate the quantity of heat released when 5.00 g of $H_2O_2(l)$ decomposes at constant pressure.

Answer: −14.4 kJ

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Sample Exercise 5.5 Relating Heat, Temperature Change, and Heat Capacity

(a) How much heat is needed to warm 250 g of water (about 1 cup) from 22 °C (about room temperature) to 98 °C (near its boiling point)? **(b)** What is the molar heat capacity of water?

Solution

Analyze In part **(a)** we must find the quantity of heat (q) needed to warm the water, given the mass of water (m), its temperature change (ΔT), and its specific heat (C_s). In part **(b)** we must calculate the molar heat capacity (heat capacity per mole, C_m) of water from its specific heat (heat capacity per gram).

Plan (a) Given C_s, m, and ΔT, we can calculate the quantity of heat, q, using Equation 5.22. **(b)** We can use the molar mass of water and dimensional analysis to convert from heat capacity per gram to heat capacity per mole.

Solve

(a) The water undergoes a temperature change of

$$\Delta T = 98\ °C - 22\ °C = 76\ °C = 76\ K$$

Using Equation 5.22, we have

$$q = C_s \times m \times \Delta T$$
$$= (4.18\ \text{J/g–K})(250\ \text{g})(76\ \text{K}) = 7.9 \times 10^4\ \text{J}$$

(b) The molar heat capacity is the heat capacity of one mole of substance. Using the atomic weights of hydrogen and oxygen, we have

$$C_m = \left(4.18\ \frac{\text{J}}{\text{g-K}} \right)\left(\frac{18.0\ \text{g}}{1\ \text{mol}} \right) = 75.2\ \text{J/mol-K}$$

$$1\ \text{mol}\ H_2O = 18.0\ \text{g}\ H_2O$$

From the specific heat given in part (a), we have

Sample Exercise 5.5 Relating Heat, Temperature Change, and Heat Capacity

Continued

Practice Exercise

(a) Large beds of rocks are used in some solar-heated homes to store heat. Assume that the specific heat of the rocks is 0.82 J/g–K. Calculate the quantity of heat absorbed by 50.0 kg of rocks if their temperature increases by 12.0 °C. **(b)** What temperature change would these rocks undergo if they emitted 450 kJ of heat?

Answer: **(a)** 4.9×105 J, **(b)** 11 K decrease = 11 °C decrease

Sample Exercise 5.6 Measuring ΔH Using a Coffee-Cup Calorimeter

When a student mixes 50 mL of 1.0 M HCl and 50 mL of 1.0 M NaOH in a coffee-cup calorimeter, the temperature of the resultant solution increases from 21.0 °C to 27.5 °C. Calculate the enthalpy change for the reaction in kj/mol HCl, assuming that the calorimeter loses only a negligible quantity of heat, that the total volume of the solution is 100 mL, that its density is 1.0 g/mL, and that its specific heat is 4.18 J/g–K.

Solution

Analyze Mixing solutions of HCl and NaOH results in an acid–base reaction:

$$HCl(aq) + NaOH(aq) \longrightarrow H_2O(l) + NaCl(aq)$$
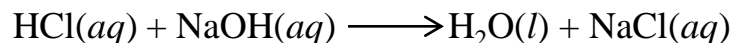
We need to calculate the heat produced per mole of HCl, given the temperature increase of the solution, the number of moles of HCl and NaOH involved, and the density and specific heat of the solution.

Plan The total heat produced can be calculated using Equation 5.23. The number of moles of HCl consumed in the reaction must be calculated from the volume and molarity of this substance, and this amount is then used to determine the heat produced per mol HCl.

Solve

Because the total volume of the solution is 100 mL, its mass is

$$(100 \text{ mL})(1.0 \text{ g/mL}) = 100 \text{ g}$$

The temperature change is

$$\Delta T = 27.5 \text{ °C} - 21.0 \text{ °C} = 6.5 \text{ °C} = 6.5 \text{ K}$$

Using Equation 5.23, we have

$$q_{rxn} = -C_s \times m \times \Delta T$$

$$= -(4.18 \text{ J/g–K})(100 \text{ g})(6.5 \text{ K}) = -2.7 \times 10^3 \text{ J} = -2.7 \text{ kJ}$$

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Continued

Because the process occurs at constant pressure,

$$\Delta H = q_P = -2.7 \text{ kJ}$$

To express the enthalpy change on a molar basis, we use the fact that the number of moles of HCl is given by the product of the volume (50mL = 0.050 L) and concentration (1.0 M = 1.0 mol/L) of the HCl solution:

$$(0.050 \text{ L})(1.0 \text{ mol/L}) = 0.050 \text{ mol}$$

Thus, the enthalpy change per mole of HCl is

$$\Delta H = -2.7 \text{ kJ}/0.050 \text{ mol} = -54 \text{ kJ/mol}$$

Check ΔH is negative (exothermic), which is expected for the reaction of an acid with a base and evidenced by the fact that the reaction causes the temperature of the solution to increase. The magnitude of the molar enthalpy change seems reasonable.

Practice Exercise

When 50.0 mL of 0.100 M AgNO3 and 50.0 mL of 0.100 M HCl are mixed in a constant-pressure calorimeter, the temperature of the mixture increases from 22.30 °C to 23.11 °C. The temperature increase is caused by the following reaction:

$$\text{AgNO3}(aq) + \text{HCl}(aq) \longrightarrow \text{AgCl}(s) + \text{HNO3}(aq)$$
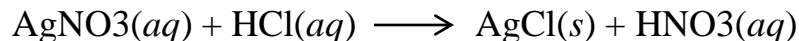
Calculate ΔH for this reaction in AgNO3, assuming that the combined solution has a mass of 100.0 g and a specific heat of 4.18 J/g °C.

Answer: −68,000 J/mol = −68 kJ/mol

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Sample Exercise 5.7 Measuring q_{rxn} Using a Bomb Calorimeter

The combustion of methylhydrazine (CH_6N_2), a liquid rocket fuel, produces $N_2(g)$, $CO_2(g)$, and $H_2O(l)$:

$$2\ CH_6N_2(l) + 5\ O_2(g) \longrightarrow 2\ N_2(g) + 2\ CO_2(g) + 6\ H_2O(l)$$

When 4.00 g of methylhydrazine is combusted in a bomb calorimeter, the temperature of the calorimeter increases from 25.00 °C to 39.50 °C. In a separate experiment the heat capacity of the calorimeter is measured to be 7.794 kJ/ °C. Calculate the heat of reaction for the combustion of a mole of CH_6N_2.

Solution

Analyze We are given a temperature change and the total heat capacity of the calorimeter. We are also given the amount of reactant combusted. Our goal is to calculate the enthalpy change per mole for combustion of the reactant.

Plan We will first calculate the heat evolved for the combustion of the 4.00-g sample. We will then convert this heat to a molar quantity.

Solve

For combustion of the 4.00-g sample of methylhydrazine, the temperature change of the calorimeter is

$$\Delta T = (39.50\ °C - 25.00\ °C) = 14.50\ °C$$

We can use ΔT and the value for C_{cal} to calculate the heat of reaction (Equation 5.24):

$$q_{rxn} = -C_{cal} \times \Delta T = -(7.794\ kJ/°C)(14.50\ °C) = -113.0\ kJ$$

We can readily convert this value to the heat of reaction for a mole of CH_6N_2:

$$\left(\frac{-113.0\ kJ}{4.00\ g\ CH_6N_2} \right) \times \left(\frac{46.1\ g\ CH_6N_2}{1\ mol\ CH_6N_2} \right) = -1.30 \times 10^3\ kJ/mol\ CH_6N_2$$

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Sample Exercise 5.7 Measuring q_{rxn} Using a Bomb Calorimeter

Continued

Check The units cancel properly, and the sign of the answer is negative as it should be for an exothermic reaction. The magnitude of the answer seems reasonable.

Practice Exercise

A 0.5865-g sample of lactic acid ($HC_3H_5O_3$) is burned in a calorimeter whose heat capacity is 4.812 kJ/°C. The temperature increases from 23.10 °C to 24.95 °C. Calculate the heat of combustion of lactic acid **(a)** per gram and **(b)** per mole.

Answer: **(a)** –15.2 kJ/g, **(b)** –1370 kJ/mol

Sample Exercise 5.8 Using Hess's Law to Calculate ΔH

The enthalpy of reaction for the combustion of C to CO_2 is -393.5 kJ/mol C, and the enthalpy for the combustion of CO to CO_2 is -283.0 kJ/mol C:

1. $\quad C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5$ kJ

2. $CO(g) + \frac{1}{2} O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -283.0$ kJ

Using these data, calculate the enthalpy for the combustion of C to CO:

3. $\quad C(s) + \frac{1}{2} O_2(g) \longrightarrow CO(g) \qquad \Delta H = ?$
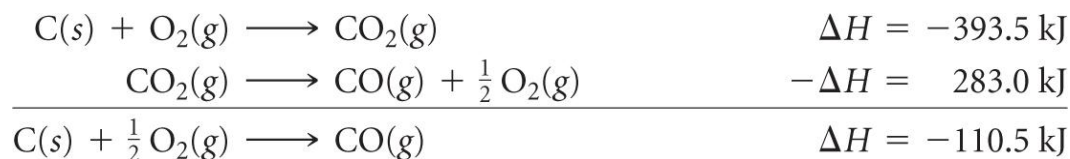
Solution

Analyze We are given two thermochemical equations, and our goal is to combine them in such a way as to obtain the third equation and its enthalpy change.

Plan We will use Hess's law. In doing so, we first note the numbers of moles of substances among the reactants and products in the target equation, (3).We then manipulate equations (1) and (2) to give the same number of moles of these substances, so that when the resulting equations are added, we obtain the target equation. At the same time, we keep track of the enthalpy changes, which we add.

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Sample Exercise 5.8 Using Hess's Law to Calculate ΔH

Continued

Solve To use equations (1) and (2), we arrange them so that C(s) is on the reactant side and CO(g) is on the product side of the arrow, as in the target reaction, equation (3). Because equation (1) has C(s) as a reactant, we can use that equation just as it is. We need to turn equation (2) around, however, so that CO(g) is a product. Remember that when reactions are turned around, the sign of ΔH is reversed. We arrange the two equations so that they can be added to give the desired equation:

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad\qquad \Delta H = -393.5 \text{ kJ}$$
$$\underline{CO_2(g) \longrightarrow CO(g) + \tfrac{1}{2}O_2(g) \qquad\qquad -\Delta H = 283.0 \text{ kJ}}$$
$$C(s) + \tfrac{1}{2}O_2(g) \longrightarrow CO(g) \qquad\qquad \Delta H = -110.5 \text{ kJ}$$

When we add the two equations, $CO_2(g)$ appears on both sides of the arrow and therefore cancels out. Likewise, ½ 0_2 (g) is eliminated from each side.

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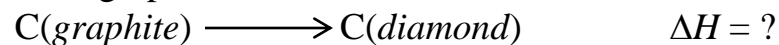# Sample Exercise 5.8 Using Hess's Law to Calculate ΔH

Continued

Practice Exercise

Carbon occurs in two forms, graphite and diamond. The enthalpy of the combustion of graphite is –393.5 kJ/mol, and that of diamond is –395.4 kJ/mol:

$$C(graphite) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5 \text{ kJ}$$
$$C(diamond) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -395.4 \text{ kJ}$$

Calculate ΔH for the conversion of graphite to diamond:

$$C(graphite) \longrightarrow C(diamond) \qquad \Delta H = ?$$

Answer: +1.9 kJ

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Sample Exercise 5.9 Using Three Equations with Hess's Law to Calculate ΔH

Calculate ΔH for the reaction

$$2\,C(s) + H_2(g) \longrightarrow C_2H_2(g)$$
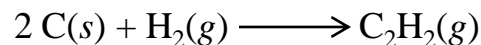
given the following chemical equations and their respective enthalpy changes:

$$C_2H_2(g) + \tfrac{5}{2}O_2(g) \longrightarrow 2\,CO_2(g) + H_2O(l) \qquad \Delta H = -1299.6\ kJ$$

$$C(s) + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H = -393.5\ kJ$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l) \qquad \Delta H = -285.8\ kJ$$
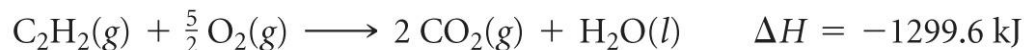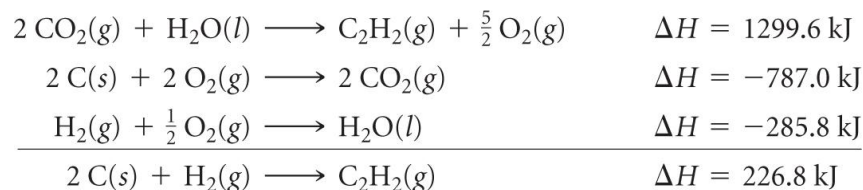
Solution

Analyze We are given a chemical equation and asked to calculate its ΔH using three chemical equations and their associated enthalpy changes.

Plan We will use Hess's law, summing the three equations or their reverses and multiplying each by an appropriate coefficient so that they add to give the net equation for the reaction of interest. At the same time, we keep track of the ΔH values, reversing their signs if the reactions are reversed and multiplying them by whatever coefficient is employed in the equation.

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Continued

Solve Because the target equation has C_2H_2 as a product, we turn the first equation around; the sign of ΔH is therefore changed. The desired equation has 2 C(*s*) as a reactant, so we multiply the second equation and its ΔH by 2. Because the target equation has H_2 as a reactant, we keep the third equation as it is. We then add the three equations and their enthalpy changes in accordance with Hess's law:

$$2\ CO_2(g)\ +\ H_2O(l)\ \longrightarrow\ C_2H_2(g)\ +\ \tfrac{5}{2}\ O_2(g) \qquad \Delta H\ =\ 1299.6\ \text{kJ}$$
$$2\ C(s)\ +\ 2\ O_2(g)\ \longrightarrow\ 2\ CO_2(g) \qquad \Delta H\ =\ -787.0\ \text{kJ}$$
$$\underline{H_2(g)\ +\ \tfrac{1}{2}\ O_2(g)\ \longrightarrow\ H_2O(l) \qquad \Delta H\ =\ -285.8\ \text{kJ}}$$
$$2\ C(s)\ +\ H_2(g)\ \longrightarrow\ C_2H_2(g) \qquad \Delta H\ =\ 226.8\ \text{kJ}$$

When the equations are added, there are 2 CO_2, 5/2 O_2, and H_2O on both sides of the arrow. These are canceled in writing the net equation.
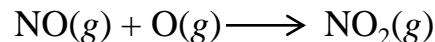
Check The procedure must be correct because we obtained the correct net equation. In cases like this you should go back over the numerical manipulations of the ΔH values to ensure that you did not make an inadvertent error with signs.

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Sample Exercise 5.9 Using Three Equations with Hess's Law to Calculate ΔH

Continued

Practice Exercise
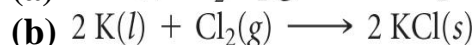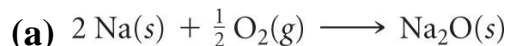
Calculate ΔH for the reaction

$$NO(g) + O(g) \longrightarrow NO_2(g)$$

given the following information

$$NO(g) + O_3(g) \longrightarrow NO_2(g) + O_2(g) \qquad \Delta H = -198.9 \text{ kJ}$$

$$O_3(g) \longrightarrow \tfrac{3}{2}O_2(g) \qquad \Delta H = -142.3 \text{ kJ}$$

$$O_2(g) \longrightarrow 2\,O(g) \qquad \Delta H = 495.0 \text{ kJ}$$

Answer: –304.1 kJ

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Sample Exercise 5.10 Equations Associated with Enthalpies of Formation

For which of these reactions at 25 °C does the enthalpy change represent a standard enthalpy of formation? For each that does not, what changes are needed to make it an equation whose ΔH is an enthalpy of formation?

(a) $2\,\text{Na}(s) + \frac{1}{2}\,\text{O}_2(g) \longrightarrow \text{Na}_2\text{O}(s)$

(b) $2\,\text{K}(l) + \text{Cl}_2(g) \longrightarrow 2\,\text{KCl}(s)$

(c) $\text{C}_6\text{H}_{12}\text{O}_6(s) \longrightarrow 6\,\text{C}(diamond) + 6\,\text{H}_2(g) + 3\,\text{O}_2(g)$

Solution

Analyze The standard enthalpy of formation is represented by a reaction in which each reactant is an element in its standard state and the product is one mole of the compound.

Plan We need to examine each equation to determine (1) whether the reaction is one in which one mole of substance is formed from the elements, and (2) whether the reactant elements are in their standard states.

Solve In (a) 1 mol Na_2O is formed from the elements sodium and oxygen in their proper states, solid Na and O_2 gas, respectively. Therefore, the enthalpy change for reaction (a) corresponds to a standard enthalpy of formation.

In (b) potassium is given as a liquid. It must be changed to the solid form, its standard state at room temperature. Furthermore, two moles of product are formed, so the enthalpy change for the reaction as written is twice the standard enthalpy of formation of $\text{KCl}(s)$. The equation for the formation reaction of 1 mol of $\text{KCl}(s)$ is

$$\text{K}(s) + \frac{1}{2}\,\text{Cl}_2(g) \longrightarrow \text{KCl}(s)$$

Continued

Reaction (c) does not form a substance from its elements. Instead, a substance decomposes to its elements, so this reaction must be reversed. Next, the element carbon is given as diamond, whereas graphite is the standard state of carbon at room temperature and 1 atm pressure. The equation that correctly represents the enthalpy of formation of glucose from its elements is

$$6\ C(graphite) + 6\ H_2(g) + 3\ O_2(g) \longrightarrow C_6H_{12}O_6(s)$$
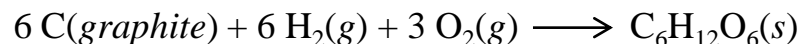
Practice Exercise

Write the equation corresponding to the standard enthalpy of formation of liquid carbon tetrachloride (CCl_4).

Answer: $C(graphite) + 2\ Cl_2(g) \longrightarrow CCl_4(l)$

Sample Exercise 5.11 Calculating an Enthalpy of Reaction from Enthalpies of Formation

(a) Calculate the standard enthalpy change for the combustion of 1 mol of benzene, $C_6H_6(l)$, to $CO_2(g)$ and $H_2O(l)$.
(b) Compare the quantity of heat produced by combustion of 1.00 g propane with that produced by 1.00 g benzene.

Solution

Analyze (a) We are given a reaction [combustion of $C_6H_6(l)$ to form $CO_2(g)$ and $H_2O(l)$] and asked to calculate its standard enthalpy change, $\Delta H°$. **(b)** We then need to compare the quantity of heat produced by combustion of 1.00 g C_6H_6 with that produced by 1.00 g C_3H_8, whose combustion was treated previously in the text. (See Equations 5.29 and 5.30.)
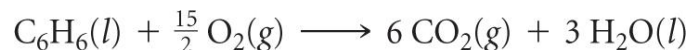
Plan (a) We need to write the balanced equation for the combustion of C_6H_6. We then look up values in Appendix C or in Table 5.3 and apply Equation 5.31 to calculate the enthalpy change for the reaction. **(b)** We use the molar mass of C_6H_6 to change the enthalpy change per mole to that per gram. We similarly use the molar mass of C_3H_8 and the enthalpy change per mole calculated in the text previously to calculate the enthalpy change per gram of that substance.

Solve

(a) We know that a combustion reaction involves $O_2(g)$ as a reactant. Thus, the balanced equation for the combustion reaction of 1 mol $C_6H_6(l)$ is

$$C_6H_6(l) + \tfrac{15}{2} O_2(g) \longrightarrow 6\, CO_2(g) + 3\, H_2O(l)$$

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Continued

We can calculate $\Delta H°$ for this reaction by using Equation 5.31 and data in Table 5.3. Remember to multiply the $\Delta H_f°$ value for each substance in the reaction by that substance's stoichiometric coefficient. Recall also that $\Delta H_f° = 0$ for any element in its most stable form under standard conditions, so $\Delta H_f°[O_2(g)] = 0$:

$$\Delta H_{rxn}° = [6\Delta H_f°(CO_2) + 3\Delta H_f°(H_2O)] - [\Delta H_f°(C_6H_6) + \tfrac{15}{2}\Delta H_f°(O_2)]$$
$$= [6(-393.5\text{ kJ}) + 3(-285.8\text{ kJ})] - [(49.0\text{ kJ}) + \tfrac{15}{2}(0\text{ kJ})]$$
$$= (-2361 - 857.4 - 49.0)\text{ kJ}$$
$$= -3267\text{ kJ}$$

(b) From the example worked in the text, for the combustion of 1 mol of propane. In part (a) of this Exercise we determined that for the combustion of 1 mol benzene. To determine the heat of combustion per gram of each substance, we use the molar masses to convert moles to grams:

$C_3H_8(g)$: $(-2220\text{ kJ>mol})(1\text{ mol>44.1 g}) = -50.3\text{ kJ>g}$

$C_6H_6(l)$: $(-3267\text{ kJ>mol})(1\text{ mol>78.1 g}) = -41.8\text{ kJ>g}$

TABLE 5.3 • Standard Enthalpies of Formation, $\Delta H_f°$, at 298 K

Substance	Formula	$\Delta H_f°$ (kJ/mol)	Substance	Formula	$\Delta H_f°$ (kJ/mol)
Acetylene	$C_2H_2(g)$	226.7	Hydrogen chloride	$HCl(g)$	−92.30
Ammonia	$NH_3(g)$	−46.19	Hydrogen fluoride	$HF(g)$	−268.60
Benzene	$C_6H_6(l)$	49.0	Hydrogen iodide	$HI(g)$	25.9
Calcium carbonate	$CaCO_3(s)$	−1207.1	Methane	$CH_4(g)$	−74.80
Calcium oxide	$CaO(s)$	−635.5	Methanol	$CH_3OH(l)$	−238.6
Carbon dioxide	$CO_2(g)$	−393.5	Propane	$C_3H_8(g)$	−103.85
Carbon monoxide	$CO(g)$	−110.5	Silver chloride	$AgCl(s)$	−127.0
Diamond	$C(s)$	1.88	Sodium bicarbonate	$NaHCO_3(s)$	−947.7
Ethane	$C_2H_6(g)$	−84.68	Sodium carbonate	$Na_2CO_3(s)$	−1130.9
Ethanol	$C_2H_5OH(l)$	−277.7	Sodium chloride	$NaCl(s)$	−410.9
Ethylene	$C_2H_4(g)$	52.30	Sucrose	$C_{12}H_{22}O_{11}(s)$	−2221
Glucose	$C_6H_{12}O_6(s)$	−1273	Water	$H_2O(l)$	−285.8
Hydrogen bromide	$HBr(g)$	−36.23	Water vapor	$H_2O(g)$	−241.8

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Comment Both propane and benzene are hydrocarbons. As a rule, the energy obtained from the combustion of a gram of hydrocarbon is between 40 and 50 kJ.

Practice Exercise

Use Table 5.3 to calculate the enthalpy change for the combustion of 1 mol of ethanol:
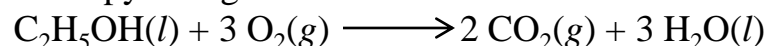
$$C_2H_5OH(l) + 3\ O_2(g) \longrightarrow 2\ CO_2(g) + 3\ H_2O(l)$$

TABLE 5.3 • Standard Enthalpies of Formation, ΔH_f°, at 298 K

Substance	Formula	ΔH_f° (kJ/mol)	Substance	Formula	ΔH_f° (kJ/mol)
Acetylene	$C_2H_2(g)$	226.7	Hydrogen chloride	$HCl(g)$	−92.30
Ammonia	$NH_3(g)$	−46.19	Hydrogen fluoride	$HF(g)$	−268.60
Benzene	$C_6H_6(l)$	49.0	Hydrogen iodide	$HI(g)$	25.9
Calcium carbonate	$CaCO_3(s)$	−1207.1	Methane	$CH_4(g)$	−74.80
Calcium oxide	$CaO(s)$	−635.5	Methanol	$CH_3OH(l)$	−238.6
Carbon dioxide	$CO_2(g)$	−393.5	Propane	$C_3H_8(g)$	−103.85
Carbon monoxide	$CO(g)$	−110.5	Silver chloride	$AgCl(s)$	−127.0
Diamond	$C(s)$	1.88	Sodium bicarbonate	$NaHCO_3(s)$	−947.7
Ethane	$C_2H_6(g)$	−84.68	Sodium carbonate	$Na_2CO_3(s)$	−1130.9
Ethanol	$C_2H_5OH(l)$	−277.7	Sodium chloride	$NaCl(s)$	−410.9
Ethylene	$C_2H_4(g)$	52.30	Sucrose	$C_{12}H_{22}O_{11}(s)$	−2221
Glucose	$C_6H_{12}O_6(s)$	−1273	Water	$H_2O(l)$	−285.8
Hydrogen bromide	$HBr(g)$	−36.23	Water vapor	$H_2O(g)$	−241.8

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Answer: −1637 kJ

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Sample Exercise 5.12 Calculating an Enthalpy of Formation from Enthalpies of Reaction

The standard enthalpy change for the reaction $CaCO_3(s) \longrightarrow CaO(s) + CO_2(g)$ is 178.1 kJ. Use Table 5.3 to calculate the standard enthalpy of formation of $CaCO_3(s)$.

Solution

Analyze Our goal is to obtain $\Delta H_f^\circ \, [CaCO_3(s)]$.

Plan We begin by writing the expression for the standard enthalpy change for the reaction:

$$\Delta H_{rxn}^\circ = \Delta H_f^\circ[CaO(s)] + \Delta H_f^\circ[CO_2(g)] - \Delta H_f^\circ[CaCO_3(s)]$$

Solve Inserting the given ΔH_{rxn}° and the known ΔH_f° values from Table 5.3 or Appendix C, we have

$$178.1 = -635.5 \text{ kJ} - 393.5 \text{ kJ} - \Delta H_f^\circ[CaCO_3(s)]$$

Solving for $\Delta H_f^\circ[CaCO_3(s)]$ gives

$$\Delta H_f^\circ[CaCO_3(s)] = -1207.1 \text{ kJ/mol}$$

Check We expect the enthalpy of formation of a stable solid such as calcium carbonate to be negative, as obtained.

TABLE 5.3 • Standard Enthalpies of Formation, ΔH_f°, at 298 K

Substance	Formula	ΔH_f° (kJ/mol)	Substance	Formula	ΔH_f° (kJ/mol)
Acetylene	$C_2H_2(g)$	226.7	Hydrogen chloride	$HCl(g)$	−92.30
Ammonia	$NH_3(g)$	−46.19	Hydrogen fluoride	$HF(g)$	−268.60
Benzene	$C_6H_6(l)$	49.0	Hydrogen iodide	$HI(g)$	25.9
Calcium carbonate	$CaCO_3(s)$	−1207.1	Methane	$CH_4(g)$	−74.80
Calcium oxide	$CaO(s)$	−635.5	Methanol	$CH_3OH(l)$	−238.6
Carbon dioxide	$CO_2(g)$	−393.5	Propane	$C_3H_8(g)$	−103.85
Carbon monoxide	$CO(g)$	−110.5	Silver chloride	$AgCl(s)$	−127.0
Diamond	$C(s)$	1.88	Sodium bicarbonate	$NaHCO_3(s)$	−947.7
Ethane	$C_2H_6(g)$	−84.68	Sodium carbonate	$Na_2CO_3(s)$	−1130.9
Ethanol	$C_2H_5OH(l)$	−277.7	Sodium chloride	$NaCl(s)$	−410.9
Ethylene	$C_2H_4(g)$	52.30	Sucrose	$C_{12}H_{22}O_{11}(s)$	−2221
Glucose	$C_6H_{12}O_6(s)$	−1273	Water	$H_2O(l)$	−285.8
Hydrogen bromide	$HBr(g)$	−36.23	Water vapor	$H_2O(g)$	−241.8

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Practice Exercise

Given the following standard enthalpy change, use the standard enthalpies of formation in Table 5.3 to calculate the standard enthalpy of formation of $CuO(s)$:
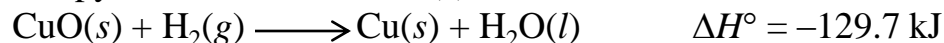
$$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l) \qquad \Delta H° = -129.7 \text{ kJ}$$

TABLE 5.3 • Standard Enthalpies of Formation, $\Delta H_f°$, at 298 K

Substance	Formula	$\Delta H_f°$ (kJ/mol)	Substance	Formula	$\Delta H_f°$ (kJ/mol)
Acetylene	$C_2H_2(g)$	226.7	Hydrogen chloride	$HCl(g)$	−92.30
Ammonia	$NH_3(g)$	−46.19	Hydrogen fluoride	$HF(g)$	−268.60
Benzene	$C_6H_6(l)$	49.0	Hydrogen iodide	$HI(g)$	25.9
Calcium carbonate	$CaCO_3(s)$	−1207.1	Methane	$CH_4(g)$	−74.80
Calcium oxide	$CaO(s)$	−635.5	Methanol	$CH_3OH(l)$	−238.6
Carbon dioxide	$CO_2(g)$	−393.5	Propane	$C_3H_8(g)$	−103.85
Carbon monoxide	$CO(g)$	−110.5	Silver chloride	$AgCl(s)$	−127.0
Diamond	$C(s)$	1.88	Sodium bicarbonate	$NaHCO_3(s)$	−947.7
Ethane	$C_2H_6(g)$	−84.68	Sodium carbonate	$Na_2CO_3(s)$	−1130.9
Ethanol	$C_2H_5OH(l)$	−277.7	Sodium chloride	$NaCl(s)$	−410.9
Ethylene	$C_2H_4(g)$	52.30	Sucrose	$C_{12}H_{22}O_{11}(s)$	−2221
Glucose	$C_6H_{12}O_6(s)$	−1273	Water	$H_2O(l)$	−285.8
Hydrogen bromide	$HBr(g)$	−36.23	Water vapor	$H_2O(g)$	−241.8

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Answer: −156.1 kJ/mol

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Sample Exercise 5.13 Comparing Fuel Values

Celery contains carbohydrates in the form of starch and cellulose, which have essentially the same fuel values when combusted in a bomb calorimeter. When we eat celery, however, our bodies receive fuel value from the starch only. What can we conclude about the difference between starch and cellulose as foods?

Solution

If cellulose does not provide fuel value, we must conclude that it is not converted in the body into CO_2 and H_2O, as starch is. A slight but critical difference in the structures of starch and cellulose explains why only starch is broken down into glucose in the body. Cellulose passes through without undergoing significant chemical change. It serves as fiber, or roughage, in the diet but provides no caloric value.

Practice Exercise

The nutrition label on a bottle of canola oil indicates that 10 g of the oil has a fuel value of 86 kcal. A similar label on a bottle of pancake syrup indicates that 60 mL (about 60 g) has a fuel value of 200 kcal. Account for the difference.

Answer: The oil has a fuel value of 8.6 kcal/g, whereas the syrup has a fuel value of about 3.3 kcal/g. The higher fuel value for the canola oil arises because the oil is essentially pure fat, whereas the syrup is a solution of sugars (carbohydrates) in water. The oil has a higher fuel value per gram; in addition, the syrup is diluted by water.

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Sample Exercise 5.14 Estimating the Fuel Value of a Food from Its Composition

(a) A 28-g (1-oz) serving of a popular breakfast cereal served with 120 mL of skim milk provides 8 g protein, 26 g carbohydrates, and 2 g fat. Using the average fuel values of these substances, estimate the fuel value (caloric content) of this serving. **(b)** A person of average weight uses about 100 Cal/mi when running or jogging. How many servings of this cereal provide the fuel value requirements to run 3 mi?

Solution

(a) Analyze The fuel value of the serving will be the sum of the fuel values of the protein, carbohydrates, and fat.

Plan We are given the masses of the protein, carbohydrates, and fat contained in a serving. We can use the data in Table 5.4 to convert these masses to their fuel values, which we can sum to get the total fuel value.

TABLE 5.4 • Compositions and Fuel Values of Some Common Foods

	Approximate Composition (% by mass)			Fuel Value	
	Carbohydrate	Fat	Protein	kJ/g	kcal/g (Cal/g)
Carbohydrate	100	—	—	17	4
Fat	—	100	—	38	9
Protein	—	—	100	17	4
Apples	13	0.5	0.4	2.5	0.59
Beer[†]	1.2	—	0.3	1.8	0.42
Bread	52	3	9	12	2.8
Cheese	4	37	28	20	4.7
Eggs	0.7	10	13	6.0	1.4
Fudge	81	11	2	18	4.4
Green beans	7.0	—	1.9	1.5	0.38
Hamburger	—	30	22	15	3.6
Milk (whole)	5.0	4.0	3.3	3.0	0.74
Peanuts	22	39	26	23	5.5

[†]Beer typically contains 3.5% ethanol, which has fuel value.

*Although fuel values represent the heat *released* in a combustion reaction, fuel values are reported as positive numbers.

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Sample Exercise 5.14 Estimating the Fuel Value of a Food from Its Composition

Continued

Solve

$$(8 \text{ g protein})\left(\frac{17 \text{ kJ}}{1 \text{ g protein}}\right) + (26 \text{ g carbohydrate})\left(\frac{17 \text{ kJ}}{1 \text{ g carbohydrate}}\right) +$$

$$(2 \text{ g fat})\left(\frac{38 \text{ kJ}}{1 \text{ g fat}}\right) = 650 \text{ kJ (to two significant figures)}$$

This corresponds to 160 kcal:

$$(650 \text{ kJ})\left(\frac{1 \text{ kcal}}{4.18 \text{ kJ}}\right) = 160 \text{ kcal}$$

Recall that the dietary Calorie is equivalent to 1kcal. Thus, the serving provides 160 Cal.

(b) Analyze Here we are faced with the reverse problem, calculating the quantity of food that provides a specific fuel value.

Plan The problem statement provides a conversion factor between Calories and miles. The answer to part (a) provides us with a conversion factor between servings and Calories.

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Sample Exercise 5.14 Estimating the Fuel Value of a Food from Its Composition

Continued

Solve We can use these factors in a straightforward dimensional analysis to determine the number of servings needed, rounded to the nearest whole number:

$$\text{Servings} = (3 \text{ mi})\left(\frac{100 \text{ Cal}}{1 \text{ mi}}\right)\left(\frac{1 \text{ serving}}{160 \text{ Cal}}\right) = 2 \text{ servings}$$

Practice Exercise

(a) Dry red beans contain 62% carbohydrate, 22% protein, and 1.5% fat. Estimate the fuel value of these beans. **(b)** During a very light activity, such as reading or watching television, the average adult expends about 7 kJ/min. How many minutes of such activity can be sustained by the energy provided by a serving of chicken noodle soup containing 13 g protein, 15 g carbohydrate, and 5 g fat?

Answer: **(a)** 15 kJ/g, **(b)** 100 min

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